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VOL. 23, SEC. B. MARCH, 1945 NUMBER 2

KINETICS OF THE CATALYTIC DEHYDROGENATION OF ALCOHOLS AND GLYCOLS IN THE GASEOUS STATE[1]

BY ARTHUR C. NEISH[2]

Abstract

The catalytic dehydrogenation of several alcohols and glycols, in the gaseous state, was found to be a first order process obeying the Arrhenius relation for variation of rate with temperature. Hence a large body of experimental data can be expressed by two constants that enable a ready comparison of different catalysts. A porous copper catalyst was found to be the most satisfactory. The fallacy of calculating the energy of activation from the temperature coefficient in this type of reaction is pointed out. Primary alcohols are dehydrogenated more slowly than secondary because the resulting aldehydes reversibly depress the activity of the catalyst, while methyl ketones have no measurable effect. Diacetyl acts as an aldehyde in this respect. The α-glycols form α-hydroxy aldehydes and ketones that depress the activity of the catalyst to a low but constant level. Rapid poisoning of the catalyst at constant temperature is a first order process. A regular relation exists between the temperature of reduction and the activity of the catalyst.

Acetoin and diacetyl can be produced in almost quantitative yields by dehydrogenation of 2,3-butanediol vapour.

Introduction

Many studies on the catalytic dehydrogenation of alcohols have been made since the pioneer work of Ipatieff (4). It is usual to report the number of millilitres of hydrogen formed per minute at a specified temperature and rate of addition of the alcohol. Sometimes the dimensions of the catalyst chamber are also given. If any of these conditions are changed, the amount of hydrogen formed per unit time will be unpredictably different. This arbitrary way of expressing the results makes it difficult to compare the activity of different catalysts, to compare the rates of dehydrogenation of different alcohols, or to study the influence of external factors on the activity of any given catalyst. Even if the apparatus is rigidly standardized, strictly comparative results are not forthcoming unless two factors are taken into account: (i) thermal expansion of the gas, which causes the time of contact to be shorter at higher temperatures, and (ii) varying degrees of dilution of the vapour by hydrogen, which also affects the time a given molecule of vapour spends in the catalyst chamber.

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The purpose of the present work is to establish a simple method for determining and expressing the activity of this type of dehydrogenation catalyst that will be independent of the actual experimental conditions.

Experimental

The rate of dehydrogenation of some common alcohols and glycols was determined using copper and nickel catalysts. The vapour was passed at a constant rate over the uniformly heated catalyst and the rate of hydrogen formation was measured after condensation of the liquid products. All experiments were done at atmospheric pressure. Isopropanol and *l*-2,3-butanediol were studied the most extensively but acetoin, ethanediol, 1,2-propanediol, ethanol, *n*-propanol, *n*-butanol, *sec.*-butanol, *n*-octanol-1, and *n*-octanol-2 were also used.

Experimental Design

The rate of dehydrogenation of isopropanol vapour, over a copper catalyst, was measured (i) at constant temperature and varying rates of flow of the isopropanol and (ii) at a constant rate of flow and varying temperatures. The data obtained were analysed to show that it was possible to express the activity of the catalyst by two constants. The work was extended to include nickel and copper–nickel catalysts and also ethanol and 2,3-butanediol.

Use was made of the condensed method of expressing the results to study the relation between the temperature of reduction and the catalytic activity of copper. The copper was oxidized by oxygen at 420° C. and reduced with hydrogen at various temperatures between 260° C. and 420° C. Its activity was then measured over a wide temperature range using isopropanol.

The rate of poisoning of the porous copper catalyst by several impure alcohols and glycols was measured at constant temperature.

The rates of dehydrogenation of all the alcohols and glycols mentioned above were compared using the porous copper catalyst. The activity of the catalyst was tested with isopropanol before and after using it for the dehydrogenation of the compound in question. The effects of adding small quantities of aldehydes or ketones to isopropanol was also studied in this manner.

Finally the dehydrogenation of *l*-2,3-butanediol was studied in some detail. The liquid products were separated and determined by fractionation through a Stedman column (18 in. long) at a reflux ratio of about 2 : 1. The material yield, space-time yield, fraction dehydrogenated, and the optical rotation of the recovered 2,3-butanediol were measured.

Analyses

The gaseous products were analysed for hydrogen, carbon dioxide, carbon monoxide, olefins, and paraffins in a Burrell gas analysis apparatus.

Acetone and octylaldehyde were determined in the presence of each other by adding an excess of isopropanol, distilling out the acetone, and estimating carbonyl groups (1) in the distillate and the residue in the distilling flask.

A Vigreux column was used. The receiver was connected to a cold trap (−80° C.) to prevent loss of acetone.

Alcohols and Glycols Used

Absolute ethanol, Merck's isopropanol (99%) and Eastman's *sec*-butanol, *n*-octanol-2 and 1,2-propanediol were used directly. The *l*-2,3,-butanediol ($[\alpha]_D^{28} = 13.10$; refractive index, 1.4318 at 26° C.) was recovered from a wheat mash after fermentation by *Aerobacillus polymyxa* (9). It was purified by fractionation through a Stedman column, as were also the *n*-butanol, *n*-octanol-1, *n*-propanol, and ethanediol used.

Preparation of Catalysts

The copper gauze catalyst consisted of rolls of 100-mesh copper gauze. It was activated by oxidation with oxygen, followed by reduction with hydrogen.

The porous copper catalyst, recommended by Hilditch (3, p. 178) was prepared by heating Merck's black copper oxide powder in a muffle furnace at 1100° C. for one hour. The sintered mass was broken up in a porcelain mortar to obtain 5- to 20-mesh particles. These were reduced with hydrogen at 320° C. No loss in volume occurred, although enough water was formed to account for reduction of 70% of the catalyst from CuO. This results in the copper being obtained in the form of porous granules.

Copper supported on magnesium oxide was obtained by adding sodium hydroxide to an equimolecular solution of copper and magnesium nitrates. The mixed hydroxides were filtered on a Büchner funnel, washed with distilled water until filtrate showed a pH of 8.0, dried at 150° C., and ground in a porcelain mortar to obtain 5- to 20-mesh particles. These were reduced with hydrogen.

The nickel– and copper–asbestos catalysts were prepared by soaking clean asbestos fibres (such as are used for Gooch crucibles) in nickel or copper nitrate, or mixtures of these, melted in their water of crystallization. The impregnated asbestos was heated at 400 to 500° C. to expel water and nitric acid. The product of this treatment (asbestos coated with the oxides) was reduced with hydrogen.

Measurement of the Available Volume in the Catalyst Chamber

The catalyst, from a known packed volume, was transferred quantitatively to a graduated cylinder containing alcohol. The volume of alcohol displaced was observed; the available volume in the catalyst chamber was obtained by subtracting this value from the total volume of the chamber. The catalysts used in this work left 78 to 85% of the volume of the chamber available for vapour.

Reactivation of the Catalysts

The catalysts were reactivated by oxidation with oxygen at 400 to 500° C. followed by reduction with hydrogen, usually at 280 to 300° C. The gases were passed into the apparatus (Fig. 1) through the stopcock at J. The rate

of flow of the hydrogen was carefully adjusted to avoid local heating. Sometimes the catalyst was reduced by an alcohol.

Measurement of the Rate of Dehydrogenation

The apparatus used is shown in Fig. 1. Current from the 6-volt battery *A* electrolysed a 10% potassium hydroxide solution in the cell *C*, the rate of electrolysis being controlled by the variable resistance *B*. The electrodes

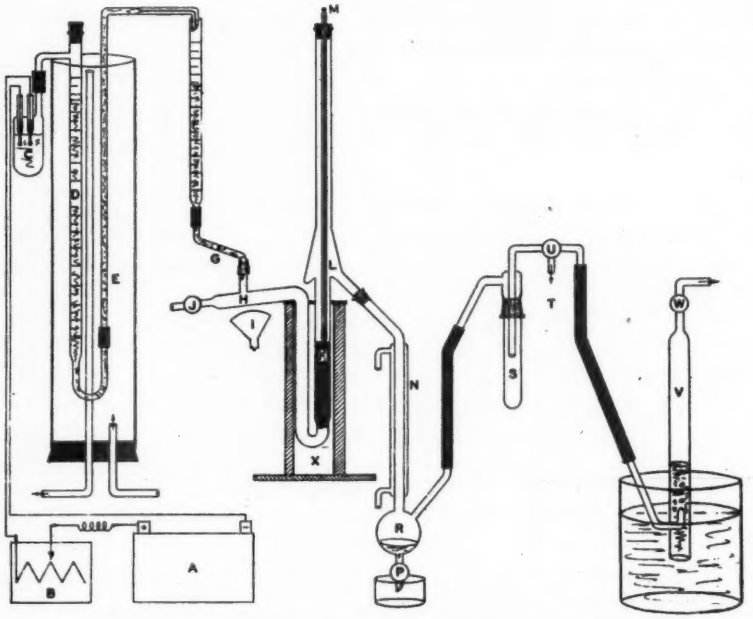

FIG. 1. *Apparatus for measuring the rate of dehydrogenation of an alcohol or glycol.*

were platinum. The gas formed in electrolysis exerted a pressure on the liquid to be dehydrogenated in the 50 cc. burette *D*, forcing it at a constant rate into the 25 cc. burette *F*, which was open to the atmosphere. The liquid flowed from *F* into the vaporizing chamber at *H* through the sintered glass disk at *G*. The rate of flow was controlled by the porosity of the disk and the height of the liquid in *F*. The rate of electrolysis was adjusted so the liquid in *F* remained at a constant level and the rate of addition of liquid was then read from *D*, which was kept at constant temperature by a current of air or water in the glass jacket *E*. This assembly served to deliver liquid into *H* at a constant, known rate. It was not satisfactory to discharge the liquid from *D* directly into *H* since this system was too sensitive to slight fluctuations in pressure that may occur in the catalyst chamber. It was only necessary to have the rate of addition constant for 15 to 20 min. at a

time, in practice, and during this period the rate of output of the liquid from *D* did not change measurably. The liquid was vaporized in the chamber *H* by means of the fish tail burner at *I*. The chamber was wrapped with wire gauze (not shown) to give uniform heating and the heat adjusted so that the liquid was vaporized as rapidly as it entered. An asbestos shield (not shown) protected *F* and *G* from the heat of the flame. The vapour passed into the part of the tube that was inside the electrically heated hot air furnace at *X*. It was preheated on passing through the descending arm of the tube and then passed through the catalyst at *K*, which was supported on a plug of glass wool in the ascending arm of the tube. The temperature of the vapour was read on the 0° to 500° C. thermometer *M* just as the vapour left the catalyst. Any condensed vapours were trapped at *L* and passed, with the vapour, into the water cooled condenser at *N*. The condensate was caught in the receiver *R* and the liquid trap *S*, which was cooled by means of a dry-ice–acetone mixture. The gas passed on through to the three-way stopcock at *U*, which could be turned so that either the gas was vented at *T* or was passed into the gas collection tube *V*. Since the volume of *V* from the two-way stopcock *W* to the tip of the gas inlet tube was accurately known, the rate of gas evolution in millilitres per minute was easily obtained. The stopcock at *W* was connected to a suction pump (not shown) so the tube *V* could be easily filled with water by opening this stopcock. The stopcock, *P*, was opened to draw off condensate, from time to time, or adjusted so the condensate was removed at the same rate as it collected.

Results

Analysis of Data

It is only necessary to know the temperature of the catalyst chamber, the rate of addition of liquid, and the rate of evolution of hydrogen under *steady* conditions for several temperatures to be able to calculate constants that describe the activity of the catalyst. The variation of the rate of hydrogen formation (space–time yield) and the degree of completion of the reaction with variation in the rate of addition of isopropanol at constant temperature, are shown in Fig. 2. These data were found to represent a first order reaction and can be reduced to a single constant, k, which is defined by the familiar expression

$$k = \frac{2.303}{t} \log \frac{1}{c} \qquad (1)$$

where t is the time of reaction and c is the fraction of unchanged isopropanol. This constant, k, obeys the Arrhenius relation for variation of rate with temperature over a wide range of experimental conditions.

Since the reaction occurs only in the catalyst chamber, the value of t is taken to be the average time a molecule of vapour spends there. It is calculated by

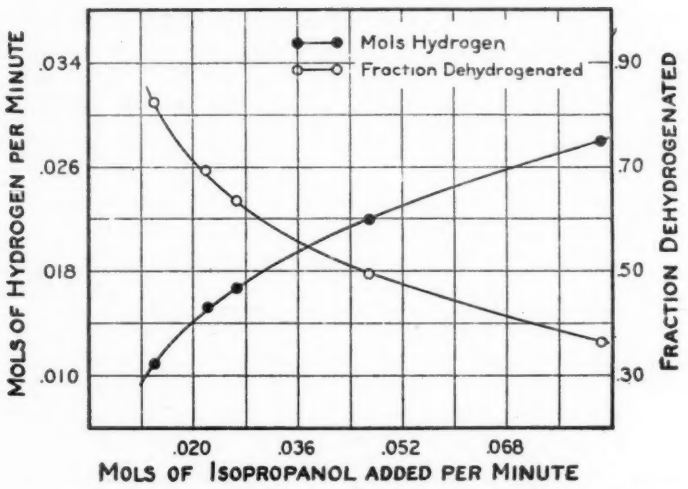

FIG. 2. *Variation of the rate of dehydrogenation of isopropanol with the rate of addition of the vapour to the catalyst chamber, at constant temperature. A copper–magnesium-oxide catalyst was used at a temperature of 299° C.*

dividing the volume available for vapour in the catalyst chamber by the volume of vapour passing through it per unit time. At a constant pressure of one atmosphere

$$t = \frac{0.7309\,V}{(N + n)T} \text{ sec.} \qquad (2)$$

N = Number of moles of isopropanol added per minute,
n = Number of moles of hydrogen formed per minute,
T = Absolute temperature,
V = Volume available for vapour in the catalyst chamber.

When t is calculated by means of Equation (2) from the data in Fig. 2 and plotted against log c a straight line is obtained (Fig. 3). This shows that the reaction is a first order process. When this is known it is possible to derive more accurate equations for the calculation of t and k from the experimental data. Equation (2) gives values of t that are 1 to 5% low, in the experimental ranges covered in this paper, since it was derived on the assumption that the volume of gas passing through the chamber per unit time is the sum of the isopropanol added and the hydrogen formed. This is not strictly true since the formation of hydrogen in the catalyst chamber does not occur instantaneously (as assumed above). Actually the composition of the gas in the catalyst chamber varies along its length as hydrogen is formed according to

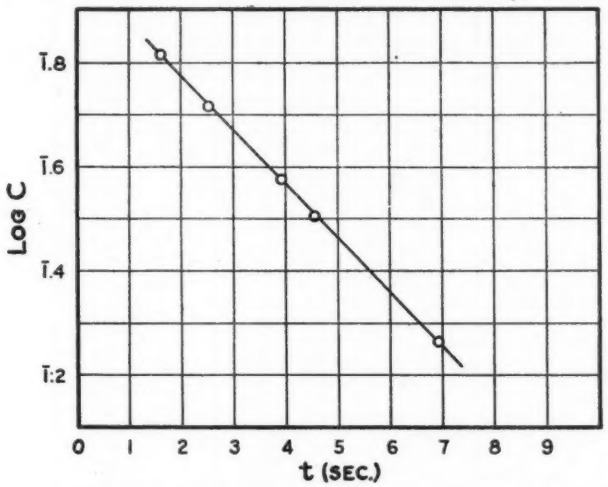

FIG. 3. *Log concentration of unchanged isopropanol plotted against time of reaction. The linear relation obtained shows the rate of the reaction to be directly proportional to the concentration of unchanged isopropanol. This comes from an analysis of the data in Fig. 2.*

the first order relation. When this is taken into account* Equation (2) becomes

$$t = \tfrac{1}{2}\left[\frac{V}{82.1\ NT} + \frac{nt}{2.303\ N \log \frac{N}{N-n}}\right] \qquad (3)$$

and Equation (1) becomes, for convenience in calculation,

$$k = \frac{164.2\ NT}{V}\left[2.303 \log \frac{N}{N-n} - \frac{n}{2N}\right] \qquad (4)$$

The dehydrogenation of ethanol and 2,3-butanediol are also first order processes (Fig. 4). However, if the extent of dehydrogenation is carried further than shown, the value of k decreases markedly in the case of 2,3-butanediol. This is not noticeable with isopropanol unless more than 80% of it is being dehydrogenated, but becomes apparent with 2,3-butanediol if the extent of dehydrogenation exceeds 45%. The experimental range is sufficiently wide to make it possible to use k as a measure of the activity of the catalyst, even in the case of 2,3-butanediol. While k is a measure of the activity of the catalyst, at a stated temperature, in order to have a complete

* *There are G gm. per min. passing every section of the tube. Hence* $G\ \frac{V_0(N+n)}{N}$ *is the flow in cubic centimetres per minute, where* V_0 *is the specific volume of the entering gas. If the gas flows the distance dx in time dt: flow in cc. per min. times dt = cross-sectional area (in sq. cm.) times dx cm. or* $G\ \frac{(N+n)}{(N)}\ V_0 dt = A\ dx$, *where A is cross-sectional area of tube. Substitution of dt from equation for first order process* $\frac{dn}{dt} = k(N-n)$, *followed by integration for whole length of the tube* ($\int A dx = V$) *leads to Equations (3) and (4) above.*

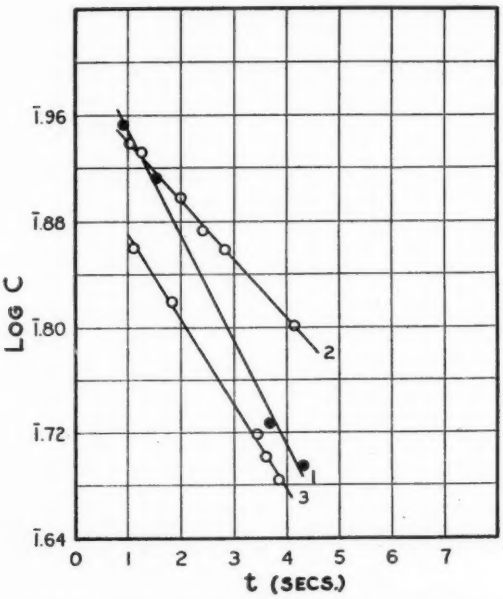

FIG. 4. *Rate of dehydrogenation of ethanol and l-2,3-butanediol at constant temperature and various rate of passages of the vapour into the catalyst chamber. Measurements were made using a porous copper catalyst. (1) Ethanol at 312° C.; (2) l-2,3-butanediol at 332° C.; (3) l-2,3-butanediol at 334° C. In each case the rate of the reaction is directly proportional to the concentration of unchanged alcohol or glycol. Lines (2) and (3) represent the catalyst at different levels of activity.*

measure of the activity of the catalyst it is necessary to know how k varies with temperature. This is given by the Arrhenius relation for variation of rate with temperature, which on integration between limits T_1 and T_2 is expressed as:

$$E = 4.576 \log \frac{k_2}{k_1} \frac{T_2 \cdot T_1}{(T_2 - T_1)} . \qquad (5)$$

When several values of k are determined at various temperatures and log k is plotted against $1/T$ a straight line is obtained if the Arrhenius relation holds true.

In more than a hundred experiments this relation has been found to hold for a variety of alcohols and dehydrogenation catalysts. The most convenient method of obtaining a value for E is to plot the data as in Fig. 5, and then read the slope of the line from the graph. It is necessary to determine both k and E to describe adequately the activity of a catalyst since one catalyst may dehydrogenate an alcohol more rapidly than another at a certain temperature, while the reverse is true at a higher temperature (Fig. 6).

Palmer (5) was able to show, without calculating t, that the Arrhenius relation was obeyed when a large excess of alcohol relative to the catalyst was

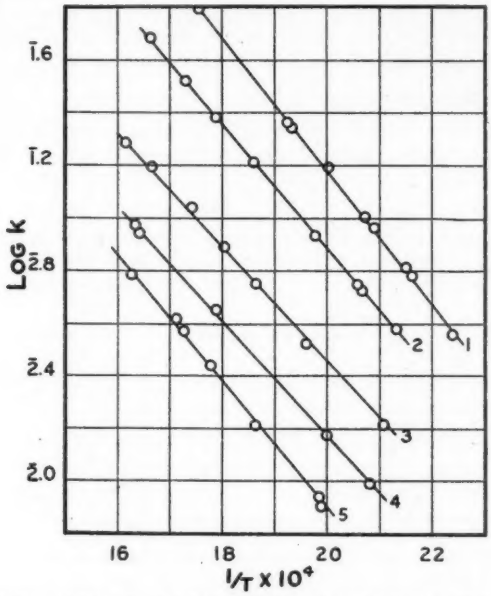

FIG. 5. *Variation of the rate of dehydrogenation of isopropanol with temperature. The Arrhenius relation is obeyed. See Table I for explanation.*

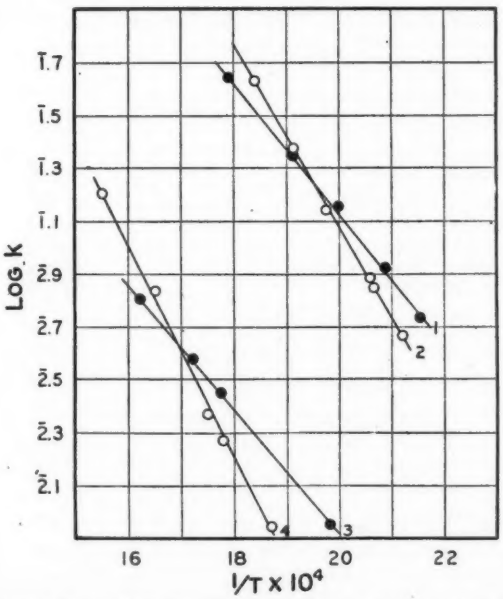

FIG. 6. *Variation of rate of dehydrogenation of isopropanol with temperature. See Table I for explanation.*

used. In his experiments t would not be appreciably affected by changing the temperature because it was small in all instances.

Variation of Catalytic Activity

The linear relation between log k and $1/T$ has been found to hold for all catalysts tested. The data in Table II demonstrate how readily catalysts

TABLE I

VARIATION OF THE RATE OF DEHYDROGENATION OF ISOPROPANOL WITH TEMPERATURE OVER CATALYSTS AT DIFFERENT LEVELS OF ACTIVITY

Line number		Catalyst	k at 253° C.	$E/10^3$
1	Fig. 5	Cu + Ni (1 : 1) on asbestos	0.3214	11.7
2		Cu on MgO	0.1585	10.8
3		Cu on MgO	0.0558	10.1
4		Cu on MgO	0.0292	9.9
5		Cu on MgO	0.0166	11.5
1	Fig. 6	Porous Cu	0.2786	10.6
2		Porous Cu	0.3699	16.7
3		Cu on MgO	0.0165	11.5
4		Cu gauze	0.0099	17.8

TABLE II

COMPARISON OF THE ACTIVITY OF SOME COPPER AND NICKEL CATALYSTS FOR DEHYDROGENATION OF ISOPROPANOL

Catalyst No.	Description of catalyst	k at 253° C.	$E/10^3$
1	100-mesh Cu gauze	0.060	10.4
2	Porous Cu		
	(*a*) Initial activity	0.179	14.1
	(*b*) After first reactivation	0.309	13.7
3	Cu on MgO		
	(*a*) Initial activity	0.042	11.8
	(*b*) After first reactivation	0.131	11.1
	(*c*) After second reactivation	0.158	10.8
4	Cu on asbestos, initial activity	0.123	15.8
5	Cu : Ni : : 4 : 1 on asbestos, initial activity	0.143	12.1
6	Cu : Ni : : 1 : 4 on asbestos, initial activity	0.406	14.7
7	Cu : Ni : : 1 : 1 on asbestos, initial activity	0.321	11.7
8	Cu : Ni : : 1 : 6 on asbestos, initial activity	0.231	10.3
9	Cu : Ni : : 6 : 1 on asbestos, initial activity	0.367	8.2
10	Ni on asbestos	0.193	11.9
	(*a*) Reduced at 320° C.		
	(*b*) Reduced at 400° C.	0.129	15.6

The same weight of copper was packed into the same volume in case of copper gauze and the porous copper catalyst. Ratios in copper–nickel catalysts are molar ratios.

can be compared when the results are expressed in terms of the constants k and E. These results show the activity of the catalyst preparation accurately, but it must be kept in mind that these activities depend largely on the way the catalyst was prepared. Such factors as thoroughness of washing and temperature of reduction exert a marked effect on the activity of the catalyst obtained. For example the copper–magnesium-oxide catalyst may not have been washed thoroughly free from sodium salts, which are known to depress the catalytic activity of copper (6); thus, on the basis of these results, it is inaccurate to conclude that copper on magnesium oxide cannot be prepared in a more active state than the porous copper. The copper–nickel catalysts are more active than copper or nickel alone, possibly owing to the fact that they are more easily reduced. These catalysts, like nickel, cause side reactions to the extent of 8 to 10% with isopropanol at 320° C., while the equally active porous copper catalyst produces only about 1% of by-products (carbon monoxide, paraffins and olefins).

The porous copper catalyst was the most satisfactory of the catalysts used in this work. A single sample has been subjected to many vapours and its activity reduced by poisoning or prolonged use, yet it has always been possible to reactivate it by oxidation followed by reduction.

Its activity varies in a regular manner with the temperature of reduction (Fig. 7) if all other factors are kept as constant as possible. The catalyst was oxidized at 420° C. with oxygen and reduced at various temperatures with hydrogen. The activity was then measured over a wide range of temperatures with isopropanol. The values of E were between 10,000 and 16,000

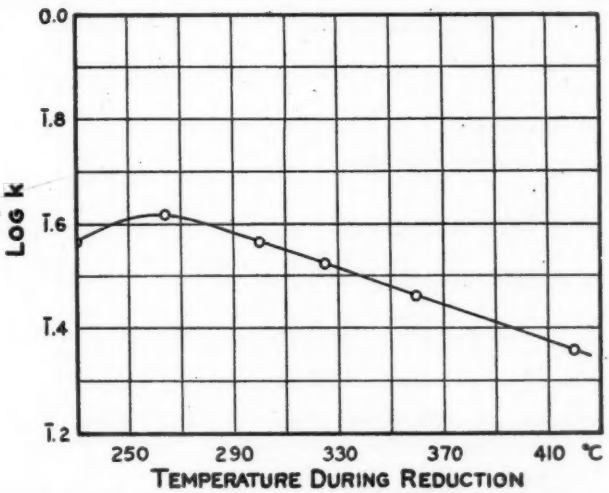

FIG. 7. *Effect of temperature of reduction on activity of the porous copper catalyst. The copper was oxidized at 420° C. with oxygen, reduced at specified temperature with hydrogen, and its activity tested with isopropanol. Log k at 253° C. is plotted against the temperature of reduction.*

cal. per mole. The time of reduction was the same in each case (three hours for 245 gm. of copper) since it is controlled by the rate of flow of the hydrogen, rather than by the temperature. When the reduction is performed in this way, log k increases linearly, as the temperature of reduction is decreased, until an optimum is reached, after which it decreases (Fig. 7). Similar results were obtained when the catalyst was oxidized by air, reduced with 2,3-butanediol, and tested on 2,3-butanediol.

These results do not agree with those of Palmer and Constable (7) who studied the reactivation of a copper catalyst by oxidation at constant temperature and reduction at various temperatures. They found a number of optima and, further, that after the catalyst had been heated above 420° C. the position of the optima shifted and they became less marked. This is in contrast to the regular results obtained above.

The activity of the freshly activated catalyst remained fairly constant for several hours with most of the alcohols tested; however, there was one notable exception, i.e., the butanol-2 used in Experiments 16 and 17 (Table IV). The catalyst showed such a rapid loss of activity that the temperature coefficient could not be measured—the change in k with time was measured at 321° C. The results are plotted in Fig. 8; log k decreases linearly with time at constant temperature and constant rate of addition of alcohol. This.

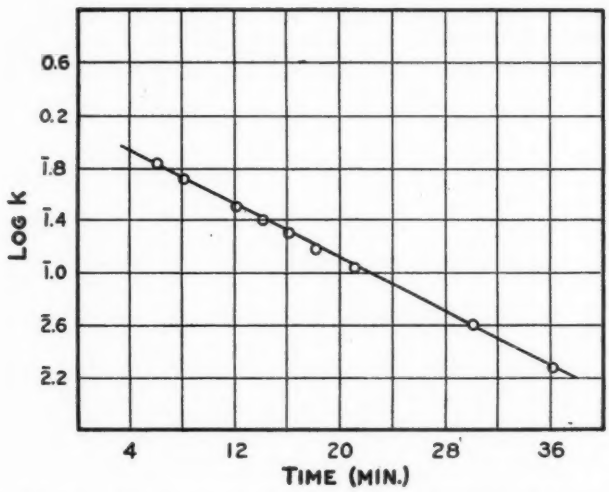

FIG. 8. *Poisoning of a copper catalyst. A sample of butanol-2 caused a rapid loss in the activity of the porous copper catalyst.. The rate of decrease of k was measured at 321° C.*

poisoning is thus a first order process, the activity being lost at a rate proportional to itself; the activity of the catalyst remains depressed when tested on isopropanol. The original level of activity is regained on activation in the usual manner.

The same type of poisoning has been obtained with impure 2,3-butanediol.

It is possible to measure the time required for the activity, k, to fall to half the original activity; typical values are given in Table III.

TABLE III

POISONING OF POROUS COPPER CATALYST BY IMPURITIES

Temperature of reaction, °C.	Liquid tested	Half-life of catalyst, hr.
321	Eastman's butanol-2	0.1
330	2,3-Butanediol (Sample 1)	3.3
320	2,3-Butanediol (Sample 2)	7.1
330	2,3-Butanediol (Sample 1 redistilled)	No loss in 10 hr.
330	Merck's isopropanol	16.0

The poisoning is caused by impurities that can be removed from the 2,3-butanediol by redistillation from a trace of sulphuric acid or sodium hydroxide. The isopropanol is sufficiently free from poisons to permit its being used to test the activity of various catalysts.

Although purified 2,3-butanediol does not show any loss in activity in 10 hr., it does not remain at a constant level of activity but passes through an optimum. This behaviour was shown whether it was redistilled from acid or alkali. Fig. 9 shows the results obtained at two different levels of activity. It is possible that slow reduction of a resistant part of the catalyst is responsible for this optimum.

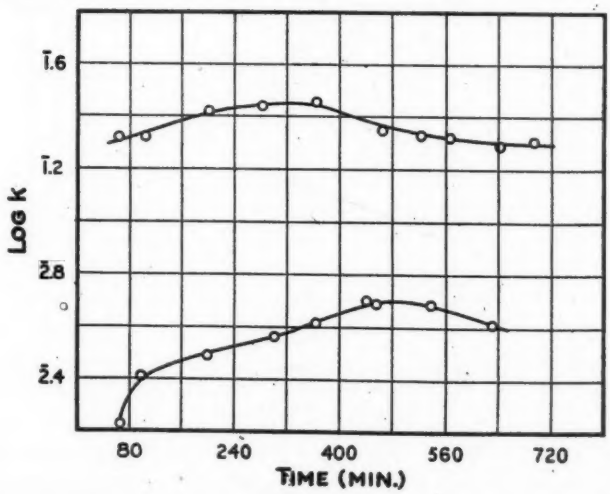

FIG. 9. *Variation of the activity of a copper catalyst with time in the absence of poisons. Tested on purified 2,3-butanediol at 330° C.*

Dehydrogenation of Various Alcohols and Glycols

The comparative rates of dehydrogenation of some common alcohols and glycols are shown in Table IV. The porous copper catalyst used was reactivated between each experiment. In most experiments the activity of the catalyst was measured with isopropanol immediately before and after use to show what change, if any, occurred in the activity of the catalyst during the measurements. Analysis of the gas showed side reactions to be negligible in each case (less than 2%). No deviations from the Arrhenius relation occurred. The value of E was calculated from the slope of the line obtained when log k was plotted against $1/T$. The examples in Fig. 10 enable one to make a quick comparison of the rates of dehydrogenation of most of the compounds tested.

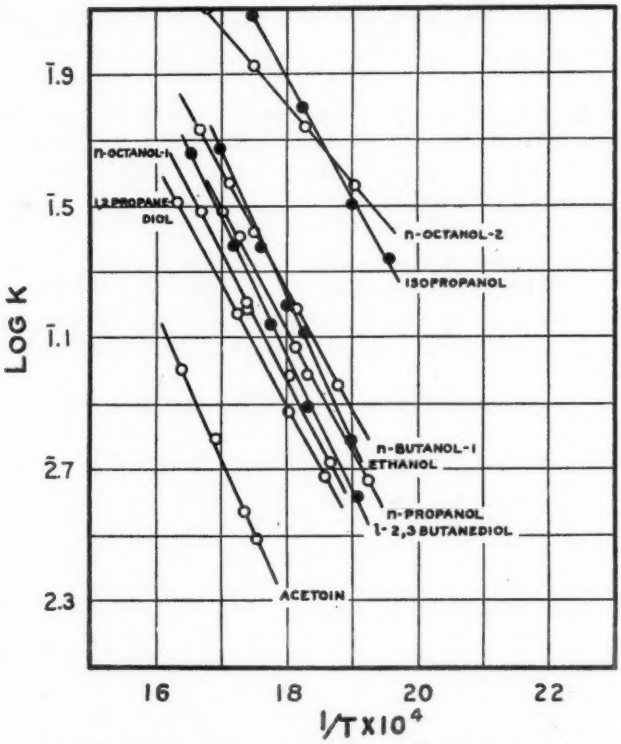

FIG. 10. *Rates of dehydrogenation of some common alcohols and glycols over a porous copper catalyst. See Table IV for more complete results. All these compounds obey the Arrhenius relation for variation of rate with temperature.*

The value of k reported from Experiment 16 (Table IV) was obtained by extrapolating the data in Fig. 8 to zero time, while the value in Experiment 17 was determined directly on butanol-2 that had been shaken with charcoal and freshly reduced copper to remove catalyst poisons. There was still

TABLE IV

COMPARATIVE RATES OF DEHYDROGENATION OF VARIOUS COMPOUNDS BY A POROUS COPPER CATALYST

Experiment No.	Compound	k at 305° C.	$E/10^3$
1	Acetoin	0.024	23.6
	Isopropanol	0.032	19.7
2	Isopropanol	0.693	8.4
	Acetoin	0.039	23.7
	Isopropanol	0.050	22.9
3	Isopropanol	1.072	7.8
	Acetoin	0.041	18.9
4	Isopropanol	0.457	8.6
	l-2,3-butanediol	0.074	13.6
5	*l*-2,3-butanediol	0.089	18.0
6	Isopropanol	0.828	7.5
	l-2,3-butanediol	0.213	20.4
7	*l*-2,3-butanediol	0.198	17.8
	Isopropanol	0.174	21.3
	l-2,3-butanediol	0.123	18.5
8	Isopropanol	0.794	6.2
	l-2,3-butanediol	0.169	16.4
	Isopropanol	0.169	16.4
9	Isopropanol	1.126	13.1
	dl-1,2-propanediol	0.134	17.9
	Isopropanol	0.215	19.8
10	Isopropanol	0.422	16.3
	Ethanediol	0.015	18.3
	Isopropanol	0.014	7.0
11	Isopropanol	1.124	12.8
	Ethanol	0.327	20.6
	Isopropanol	1.124	18.1
12	Isopropanol	0.644	7.4
	n-propanol	0.241	19.2
	Isopropanol	0.462	12.6
13	Isopropanol	1.123	12.3
	n-butanol-1	0.274	17.0
	Isopropanol	0.998	17.5
14	Isopropanol	1.146	7.0
	n-octanol-1	0.175	18.1
	Isopropanol	0.631	16.6
15	Isopropanol	0.782	9.5
	n-octanol-2	0.953	10.7
	Isopropanol	1.125	13.1
16*	Isopropanol	1.167	—
	n-butanol-2	1.160	—
17**	Isopropanol	0.640	—
	n-butanol-2	0.773	—
18	Isopropanol	0.721	17.7
	Isopropanol	0.676	18.8
	Isopropanol	0.676	18.8

* *k at 321° C. in this experiment.*
** *k at 327° C. in this experiment.*

sufficient poison remaining to make the determination of E difficult, largely because the poison seems to be more strongly adsorbed at lower temperatures.

The three glycols tested all lower the activity of the catalyst to a new level; the decrease takes place in a few minutes yet the activity may remain constant at this low level for a matter of hours. The original activity is not recovered on changing back to isopropanol but is regained on reactivation. Ethanediol causes the most marked depression, while 2,3-butanediol causes the least. It is probable that the α-hydroxyaldehydes or ketones formed as dehydrogenation products are responsible for this effect, since acetoin shows the same effect even more markedly than ethanediol. This is quite different from a poisoning effect in which the loss of activity is continuous.

The primary alcohols appear to be dehydrogenated less readily than the secondary. This agrees with the observation of Palmer and Constable (8) that isopropanol is dehydrogenated about five times as rapidly as ethanol, n-propanol, or n-butanol.

The primary alcohols have a higher value of E and a lower value of k than the secondary alcohols when compared on the same catalyst. One might be tempted to conclude from these facts that the primary alcohol group required more energy for activation than the secondary alcohol group. This is misleading, however, since in reality the primary group is dehydrogenated just about as readily as the secondary. The differences observed are caused by reversible combination of the products of dehydrogenation with the catalyst. The aldehydes show a marked tendency to lower the activity of the catalyst as does also diacetyl, while the simple methyl ketones have little or no effect. This is shown by the data in Table V. Addition of diacetyl or an aldehyde to isopropanol increases the value of E and lowers the value of k so the kinetics of dehydrogenation of isopropanol becomes similar to that of the primary alcohols. (Due allowance was made for the diluting effect of the added substance in calculating the values in Table V.) Addition of the methyl ketones, on the other hand, is without noticeable effect.

These results are best explained by assuming that the carbonyl compounds can combine with the active centres of the copper catalyst in a reversible manner, i.e., active copper + carbonyl compound $\rightleftarrows$ catalytically inactive complex. It is assumed that the equilibrium is shifted to the right on lowering the temperature, in order to account for the large temperature coefficient. In the case of diacetyl there is visual evidence for this assumption. The isopropanol–diacetyl mixture (Experiment 1, Table V) had a marked yellow colour which was all removed from the first 10 ml. of the mixture to pass over the catalyst; after that the condensate was yellow. On lowering the temperature, further adsorption of colour was noticed, while on raising it desorption occurred. These observations and the unusually high temperature coefficient are facts that support the assumption that diacetyl combines reversibly with the catalyst, the equilibrium being shifted in favour of the combined state as the temperature is lowered. This assumption can be extended to

TABLE V

EFFECT OF PRODUCTS OF DEHYDROGENATION ON THE ACTIVITY OF THE CATALYST

Experiment No.	Mixture used to test activity of catalyst	k at 305° C.	$E/10^3$
1	Isopropanol	0.676	18.8
	Isopropanol + diacetyl	0.177	31.4
	Isopropanol	0.467	17.0
2	Isopropanol	1.112	19.0
	Isopropanol + acetone	1.112	19.0
	Isopropanol	1.112	19.0
3	Isopropanol	0.761	15.9
	Isopropanol + methyl ethyl ketone	0.669	20.5
	Isopropanol	0.669	20.5
4	Isopropanol	1.112	19.0
	Isopropanol + acetaldehyde	0.391	22.9
	Isopropanol	0.761	15.9
5	Isopropanol	0.669	20.5
	Isopropanol + propionaldehyde	0.248	25.7
	Isopropanol	0.669	20.5

Solutions contained 0.04, 0.11, 0.12, 0.13, and 0.13 mole fraction of diacetyl, acetaldehyde, acetone, methyl ethyl ketone, and propionaldehyde, respectively.

explain the effect of aldehydes. The methyl ketones have such a slight tendency to combine with the catalyst they do not exert any measurable effect (Table V). Furthermore, it was shown by a direct measurement that the primary alcohol group is dehydrogenated about as readily as the secondary by dehydrogenating an equimolecular mixture of isopropanol and *n*-octanol-1 (porous copper catalyst at 325° C.) and measuring the relative amounts of acetone and octylaldehyde produced. The isopropanol was dehydrogenated only 1.4 times as rapidly, while if the two alcohols are tested separately it is dehydrogenated 5.4 times as rapidly.

The dehydrogenation of *l*-2,3-butanediol was studied more extensively than that of the other compounds (see Table VI). The products are acetoin and diacetyl, the former being identified by its boiling point, refractive index, formation of a crystalline dimer on treating with zinc dust, and by the melting point of its semicarbazone, while the latter was identified by its boiling point, refractive index, colour, and the melting point of its *bis* phenylhydrazone.

The relative amounts of these products formed from *l*-2,3-butanediol by passage of the vapour over the porous copper catalyst under various conditions are shown in Table VI. Experiments 1, 2, and 3 were made on the catalyst at the same level of activity at three different temperatures so that if log k is plotted against $1/T$ a straight line is obtained. It will be noticed that relatively more diacetyl is formed than acetoin as the temperature of the reaction is increased. This relation is shown in Fig. 11, for these three experiments. The other experiments are not quantitatively comparable to

these since the catalyst was at a different level of activity in each case. However, they agree qualitatively, for example, Experiment 5, which was

TABLE VI

DEHYDROGENATION OF *l*-2,3-BUTANEDIOL OVER A POROUS COPPER CATALYST (SINGLE PASS)

Experiment No.	Temp., °C.	*k*	Space-time* yield	*t*† sec.	Per cent of theory				Fraction dehydrogenated	Ratio‡ of acetoin to diacetyl	Optical§ rotation of recovered 2,3-butanediol
					Diacetyl	Acetoin	Recovered 2,3-butanediol	Yield of reaction			
1	291	0.144	0.0164	3.52	8.3	67.3	22.6	98.2	0.419	8.3	− 4.1
2	335	0.282	0.0225	2.81	11.6	71.6	14.1	97.3	0.474	6.3	− 3.8
3	363	0.522	0.0312	2.11	19.6	56.4	16.8	92.8	0.478	2.9	− 7.9
4	323	0.045	0.0064	8.88	11.8	68.0	19.8	99.6	0.458	5.9	− 5.9
5	382	0.059	0.0054	8.72	15.7	40.2	43.7	99.6	0.358	2.6	−11.7
6	315	0.164	0.0218	2.11	3.4	47.9	48.6	99.9	0.273	14.4	−11.6
7	300	0.045	0.0058	8.31	7.3	48.2	42.2	98.8	0.314	6.6	−10.4

* *Moles of hydrogen formed per minute per 100 cc. of catalyst space.*

† *Time of reaction (i.e., average time a molecule of the vapour spends in the catalyst chamber).*

‡ *Grams of acetoin formed divided by grams of diacetyl formed.*

§ $[\alpha]_D$ *at 28° C.; the specific optical rotation of the 2,3-butanediol used in these experiments was −13.1 and the refractive index 1.4318 at 26° C.*

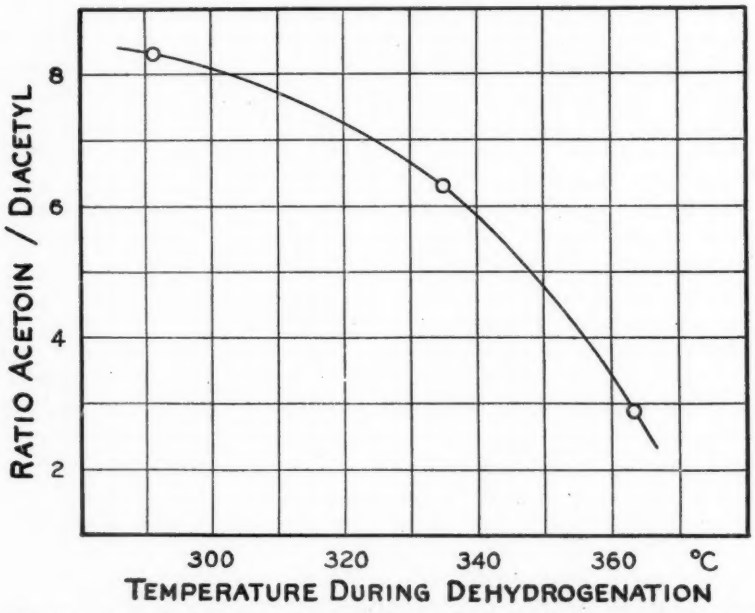

FIG. 11. *Shift of equilibrium in favour of diacetyl formation with increase in temperature. The liquid products from the dehydrogenation of l-2,3-butanediol, over a porous copper catalyst, were fractionated to determine relative amounts of acetoin and diacetyl.*

made at the highest temperature, shows the greatest proportion of diacetyl relative to acetoin, although the extent of dehydrogenation is less than in the first four experiments. Furthermore, Experiment 4, which was done at a relatively low temperature, shows a lower proportion of diacetyl, although the reaction was closer to completion than in Experiment 5.

From a consideration of all the experiments reported in Table VI, it seems that the proportion of diacetyl to acetoin is increased by (*a*) raising the temperature and (*b*) forcing the reaction closer to completion, and of these two factors the first exerts the greater effect, probably because the equilibrium shifts in favour of diacetyl formation as the temperature is increased. This explanation is supported by the changed optical rotation of the recovered 2,3-butanediol. Although the recovered butanediol has the correct refractive index the optical rotation is greatly lowered (Table VI). This can be explained by the reversible nature of the reaction.

$$\text{2,3-Butanediol} \rightleftarrows \text{acetoin} \rightleftarrows \text{diacetyl} + \text{hydrogen}$$
$$+$$
$$\text{hydrogen}$$

The 2,3-butanediol formed by hydrogenation of acetoin would be a mixture of the three possible stereoisomers; hence the lowering of the optical rotation is a measure of the extent of the reversal of the reaction. In other words, a low optical rotation signifies a close approach to equilibrium. It is apparent, then, that in Experiment 3 equilibrium was not approached as closely as in Experiments 1 and 2, probably because of the shift of the equilibrium in favour of diacetyl formation. This effect is even more apparent in Experiment 5 where the temperature was higher and the shift in favour of diacetyl greater, while the extent of the reaction was less.

The extent of the reaction, which is expressed as the fraction of the total number of moles of removable hydrogen that are removed, can be calculated either from the liquids formed or from the rate of gas evolution under steady conditions. These two methods of calculation gave the same result, thus providing a check on the accuracy of fractionation of the liquids.

The catalyst showed no appreciable change in activity during the course of the experiments (two to six hours) except in Experiment 5, where a 10% loss occurred. The material yields (Table VI) are quite high, being almost quantitative but tend to decrease as the reaction is carried closer to completion. An experiment using a nickel–asbestos catalyst gave a 92% yield at 340° C. (19% diacetyl, 47% acetoin, 26% recovered, 2,3-butanediol). This is a lower yield than that in Experiment 5, which was carried towards completion to about the same extent and shows nickel to be an undesirable catalyst.

General Discussion

The practical unit for reporting the activity of a catalyst has usually been the space–time yield (i.e., the amount of product formed per unit volume of catalyst space per unit time). In the dehydrogenation of alcohols, it varies

with the rate of passage of the vapour into the catalyst chamber as shown in Fig. 2, and also with the temperature. Thus it is not possible to calculate what the space–time yield would be under any other conditions than those for which it is reported. However, if the activity of the catalyst is reported as the constants k (at any specified temperature) and E, the space–time yield can be calculated for any desired temperature or rate of passage of vapour, in the experimental range, for a catalyst chamber of any desired size by substitution into Equation (4). Calculation of the space–time yield in this manner not only enables one to see what it is under varied conditions but also tells what the extent of the reaction is, since the fraction dehydrogenated is obtained. This is of practical importance since in such reactions the most economical rate of passage of the vapour, at any given temperature, will be determined by the increased space–time yields with increased flow rates, on one hand, and the decrease in the extent of the reaction, which occurs simultaneously, on the other hand.

The constants, k and E, are independent of the size of the apparatus. The same values were obtained for a given catalyst by using it in different chambers, one of which had six times the volume of the other. Because of their independence of the experimental conditions, the manner in which they permit the condensed expression of a large body of experimental data, and the basis they afford for further calculation, these constants form a convenient means of describing the activity of a catalyst.

On simple theoretical grounds one might expect the catalyst to act by decreasing the energy of activation. Constable (2) studied the variation of E with the activity of a copper catalyst (reduced at different temperatures) and found some tendency to relationship. In the present work all attempts to show such a relation have failed. The activity of the catalyst (as measured by k) was decreased by constant use, by sintering, and by raising the temperature of reduction, yet the value of E varied in an irregular manner. The only indication of such a relation was obtained by comparing primary and secondary alcohols on the same catalyst. In general, primary alcohols show a lower value of k and a higher value of E than secondary alcohols. However, it was shown that here, at least, E is not a true measure of the activation energy, since the higher temperature coefficient is due to reversible adsorption of the products of the reaction. Actually, primary alcohols are dehydrogenated about as readily as secondary alcohols. The lower k values are caused by a reversible depression of the activity of the catalyst by the aldehydes formed. Methyl ketones do not have this effect. Although E is not a true measure of the energy of activation in this type of reaction, it always describes the variation of k with temperature accurately.

Four distinct ways in which the activity of a copper dehydrogenation catalyst may be lowered were encountered in this work.

(1) The "natural" decay of the catalyst, which occurs slowly and is probably due to resin formation and sintering of the catalyst.

(2) The irreversible depression of the activity, which is caused by a product of the reaction, a phenomenon found in the experiments with the glycols tested. The activity is lowered to a certain level in a few minutes, yet remains constant at that level for hours.

(3) A reversible lowering of the activity caused by a product that combines reversibly with the catalyst (i.e., diacetyl or aldehydes).

(4) A rapid irreversible poisoning, which is a first order process, the activity being lost at a rate proportional to the activity at that time.

The catalyst was always restored to its original activity by oxidation, followed by reduction.

Acknowledgment

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A SIMPLE METHOD FOR DETERMINING THE COPPER CONTENT OF MILK POWDER[1]

By M. Boulet[2] and W. D. McFarlane[3]

Abstract

The procedure involves extraction of the copper by refluxing the milk powder with a solution of potassium thiocyanate in 95% acetone and determining the copper content of the extract by colorimetry with dihydroxyethyldithiocarbamic acid, after removing carotenoid pigments. The results agree with those obtained by the wet-digestion method but are higher than those yielded by dry-ashing. The advantages of the extraction method are (*a*) simplicity, (*b*) low copper content of the "blank", which, with proper precautions, may be negligible, and (*c*) a large enough sample can be used to permit visual colour comparison with standard solutions of Alizarin Yellow G.

Introduction

Copper contamination of milk powder has a marked deleterious effect on its keeping quality, and varying amounts of copper may be introduced, depending on the conditions of production, transportation, and processing of the milk. The dry milk industry, therefore, has need of a simple method for the routine estimation of copper in plant control operations.

The procedures now commonly used (1, 5, 6, 7) employ ashing or wet-digestion to destroy organic matter. Dry-ashing may cause loss of copper (4), especially with milk that has a high content of calcium phosphate. Wet-digestion requires that a relatively small sample be used, hence, a highly sensitive instrument is required for the final colour measurement. Furthermore, the "reagent blank" usually has a high copper content.

These methods are detailed and time consuming and are unsuitable for routine analysis. The procedure described below has advantages in simplicity and in the fact that a large enough sample may be used to permit visual colour comparison with standard solutions of Alizarin Yellow G.

Experimental

The writers' aim was to extract the copper directly from the milk powder with an organic solvent containing the carbamate reagent—sodium diethyldithiocarbamate. It was assumed that the carbamate reagent would combine with the copper in the milk powder, and the golden-brown coloured copper carbamate would be extracted by solvents such as acetone, carbon tetrachloride, or isoamyl alcohol. Proteins would be practically insoluble in these solvents; in the presence of the solvent they should not adsorb the copper

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carbamate complex and they should not form emulsions with the solvent when only a small amount of water was present. Furthermore, the milk fat would be dissolved and so permit better contact between the copper reagent and lipoprotein.

When 5 gm. of milk powder was shaken with 25 ml. of isoamyl alcohol (saturated with a 1% aqueous solution of the carbamate reagent) for seven hours and the colour of the extract measured at 30 min. intervals, it was found that the colour intensity increased out of proportion to the amount of copper present. Apparently the long time required to extract the copper affected the stability of the coloured copper carbamate. Attempts to reduce the extraction time by refluxing for 30 min. or to stabilize the colour by adding hydrogen peroxide, formaldehyde, or sodium cyanide, also were unsuccessful.

It was observed, however, that the rate of extraction of copper was higher when sodium cyanide was present and this was accompanied by a change in the physical nature of the milk powder, the sodium cyanide acting as a swelling or hydrating agent. Since the effect of anions on the hydration of proteins is most marked in the case of CNS^- (3, p. 268) and since potassium thiocyanate is very soluble in acetone, it appeared that a solution of potassium thiocyanate in acetone might be the ideal reagent for the extraction of copper from milk powder.

Subsequent experiments showed that the copper content of milk powder could be determined by a procedure comprised of the following steps: (*a*) extraction of the copper by refluxing the powder with a solution of potassium thiocyanate in 95% acetone, (*b*) separation of the clear extract by decantation and removal of the acetone in the extract by distillation, (*c*) suspension of the acetone soluble residue in ammonium citrate solution (citrate prevents the interference of iron in the colour reaction), (*d*) formation of the copper salt by addition of a "modified carbamate reagent," (*e*) extraction of carotenoid pigments with petroleum ether, (*f*) extraction of the coloured copper salt from the aqueous phase with isoamyl alcohol and colorimetric estimation of the copper present. In Step (*d*) it was necessary to form the copper salt prior to extracting the carotenoid pigments with petroleum ether, as part of the copper (presumably in lipoid-combination) was soluble in petroleum ether.

The so-called "modified carbamate reagent" is dihydroxyethyldithiocarbamic acid, which was recommended by Geiger and Müller (2) for the colorimetric determination of copper in aqueous solutions, the copper salt being more soluble in water than the copper salt of sodium diethyldithiocarbamate. We have found that the copper salt of dihydroxyethyldithiocarbamic acid is not extracted from aqueous solution by petroleum ether but is soluble in isoamyl alcohol; the copper salt of sodium diethyldithiocarbamate is extracted by petroleum ether and by isoamyl alcohol. Both reagents give identical colours (maximum absorption at about 435 mμ) of the same intensity with a given amount of copper and the reactions show the same deviation from Beer's law.

The Beer's law constant, K, as calculated from the relation $K = \frac{2 - \log G}{C}$, where C is the concentration of copper in micrograms per millilitre and G is the corrected galvanometer reading, was found to be 0.238 ± 0.017, with a maximum deviation of 0.034.

Description of the Method

Reagents

1. *Extracting solution.* Potassium thiocyanate (5 gm.) was dissolved in 100 ml. 95% acetone. The potassium thiocyanate must be relatively free from copper. Baker's analysed c.p. reagent was satisfactory.
2. *Modified carbamate reagent.* A 2% solution of dihydroxyethyldithiocarbamic acid was prepared by dissolving 0.2 gm. of carbon disulphide in 100 ml. of water, with the aid of vigorous shaking, and adding 2 gm. of diethanolamine.
3. *Purified ammonium citrate solution.* Citric acid (50 gm.) was dissolved in 330 ml. water and 70 ml. of concentrated ammonium hydroxide was added. To this solution in a separatory funnel was added 2 ml. of the modified carbamate reagent, and the coloured copper complex was extracted with 15-m. portions of isoamyl alcohol. Usually four extractions were sufficient.
4. *Redistilled water.* Distilled water was redistilled in an all-glass still. This redistilled water was used to prepare the 95% acetone and the modified carbamate reagent.

Procedure

A. Extraction. Milk powder (2.5 gm.) was weighed into a 125 ml. Erlenmeyer flask, 30 ml. extracting solution was added, and the mixture refluxed for 10 min. on a water-bath. The clear extract was decanted into a 125 ml. Erlenmeyer flask, 20 ml. extracting solution was added to the residue, and extraction by refluxing repeated as before. The residue was finally washed twice with 10 ml. portions of extracting solution. (When 5 gm. of milk powder was taken for analysis or when the sample had a very high copper content, as many as five successive washings were required to extract the copper completely. Absence of colour in the extract was found to be a reliable criterion of the completeness of extraction.) The extracts and washings were combined and the acetone was distilled-off. The residue was suspended in 15 ml. of ammonium citrate solution and transferred to a 125 ml. separatory funnel. The flask was rinsed with 15 ml. of ammonium citrate solution and this was added to the contents of the separatory funnel.

B. Colour reaction. The modified carbamate reagent (2 ml.) was added and, after mixing, the suspension was twice extracted with 10 ml. portions of petroleum ether. Extraction of the fat left a clear aqueous solution from which the copper salt was extracted by shaking twice with 5 ml. aliquots of isoamyl alcohol. The aqueous phase was discarded and the alcohol extracts were combined and clarified by the further addition of 5 ml. of isoamyl alcohol.

C. Colorimetry. The colour intensity was measured in a Coleman Spectrophotometer with the instrument set at 435 mμ and adjusted to 100% light transmission with pure isoamyl alcohol in the colorimeter tube. The copper content of the reagents was determined by a "blank" estimation in which the entire procedure was carried through without milk powder and starting with 50 ml. of extracting solution. The copper content of the "test" and "blank" solutions was determined by reference to a calibration curve prepared from colorimeter readings on isoamyl alcohol solutions of the copper salt containing 0.1 to 3.0 μgm. of copper per litre.

Results

The results of copper determinations on six different whole milk powders by this extraction method and by dry-ashing (6) are given in Table I. One of the powders was also analysed by the wet-digestion method (1).

TABLE I

COPPER CONTENT OF WHOLE MILK POWDERS AS DETERMINED BY DIFFERENT METHODS

Sample No.	Copper, μgm. per gm. of milk powder		
	Extraction	Dry ashing	Wet digestion
1 (*a*)	1.2	0.8	1.4
(*b*)	1.2	0.5	1.2
(*c*)	1.2	0.7	1.2
2	0.9	0.6	—
3	0.9	0.7	—
4	2.0	1.8	—
5	2.8	2.2	—
6	10.4	10.8	—

Reproducible results were obtained by the extraction method and they agreed closely with those obtained by the wet-digestion method, but the results by dry-ashing were generally lower and more variable. Known amounts of copper were added to portions of the same milk powder and the samples analysed by the extraction method with the results shown in Table II, which indicate satisfactory recoveries.

TABLE II

RECOVERY OF ADDED COPPER BY THE EXTRACTION METHOD

Sample No.	Copper, μgm. per gm. of milk powder		
	Added	Determined	Recovery, %
1	Nil	1.2	—
2	2.0	3.1	97.0
3	6.0	7.2	100.0
4	14.0	14.4	94.8
5	20.0	20.0	94.5

Modified Method Employing Visual Colorimetry

Since many "control" laboratories in powdered milk plants may not be equipped with a photoelectric colorimeter, it was thought advisable to modify the method to permit visual colour comparison with a series of standards. The intensity of the colour of copper carbamate in isoamyl alcohol increases slowly on standing, and hence these solutions are unsuitable for permanent standards. It was found, however, that the colour shade of a dilute aqueous solution of Alizarin Yellow G was indistinguishable from the colour of copper carbamate in isoamyl alcohol and the colour was quite stable. The colour standards were prepared as follows:

A sample of Alizarin Yellow G (British Drug Houses) was purified by dissolving in absolute ethyl alcohol, filtering, and evaporating the solution to dryness under reduced pressure. The powdered alcohol-soluble Alizarin Yellow G (0.1103 gm.) was dissolved in redistilled water, with the aid of heat, and the solution diluted to 100 ml. in a volumetric flask. A 1 ml. aliquot of this stock solution was transferred to a 100 ml. volumetric flask and diluted to volume with redistilled water. The colorimeter reading on this standard solution was equivalent to a copper carbamate solution containing 1 p.p.m. of copper.

A series of colour standards was prepared by selecting 10 matched test-tubes (150 by 18 mm.). Aliquots of the standard Alizarin solution, 1 to 10 ml., were pipetted into the test-tubes and each was diluted to 10 ml. with redistilled water, employing a Mohr type pipette. The contents of each tube was mixed with a swirling motion, the tubes were stoppered, and placed in order in a test-tube rack backed by a sheet of white paper or a piece of white frosted glass.

A 5 gm. sample of milk powder was taken for analysis and the extraction of copper was carried out in the manner described above. The final isoamyl alcohol solution of the coloured copper salt was transferred to a test-tube of the same dimensions as those used to prepare the colour standards.

The colour matching was carried out in daylight, the operator (standing with his back to the window and holding the test-tube rack level with the eye) fitted the sample solution into the standard series until the position of closest match was obtained. Assuming that the sample tube matched tube No. 8 in the series (8 ml. of standard Alizarin solution plus 2 ml. of water) the copper content of the milk powder was; $8/10 \times 15/5 = 2.4$ p.p.m.

The colour standards covered a range of 0.3 to 3 p.p.m. of copper in the milk powder and in intervals of 0.3 p.p.m., when a 5 gm. sample of milk powder was taken for analysis. The sample tube could be matched between a pair of standard tubes so that the error should not exceed 10%.

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LE DOSAGE POLAROGRAPHIQUE DE L'OXYGÈNE DISSOUS DANS L'EAU DE MER[1]

PAR PAUL A. GIGUÈRE[2] ET LOUIS LAUZIER[3]

Résumé

La méthode polarographique peut être employée pour déterminer rapidement et avec précision la teneur en oxygène de l'eau de mer. Des mesures faites avec l'électrode à gouttes de mercure et avec une micro-électrode de platine, soit fixe soit tournante, ont montré que le courant de diffusion est une fonction linéaire de la concentration de l'oxygène. Les sels dissous dans l'eau de mer agissent comme électrolyte fort, "supporting electrolyte", et sont en concentration suffisante pour supprimer tout maximum. En tenant compte du courant résiduel, la limite de sensibilité de cette méthode est de 0.02 cc. par litre et la précision est environ le double de celle de l'analyse chimique par la méthode de Winkler, sauf dans le cas de l'électrode de platine rotative qui a une précision beaucoup plus grande.

Introduction

L'électrolyse avec l'électrode à gouttes de mercure a été appliquée par de nombreux auteurs à la détermination de l'oxygène contenu dans l'eau. Heyrovský (1), le premier, a étudié ce phénomène et a prouvé que l'oxygène était d'abord réduit en peroxyde d'hydrogène puis en ion hydroxyle. A un pH de 7 la première réaction commence à un potentiel de -0.4 v. et la deuxième, à -1.1 v. (3, p. 295). Heyrovský (2), Vitek (4, p. 311), de même que Kolthoff et Miller (5) ont étudié certains aspects du problème, tels que l'analyse de l'oxygène de l'air et l'influence du pH sur le potentiel de polarisation de l'oxygène. Enfin Manning (7) a étendu cette méthode à l'étude de l'eau des lacs. Elle ne semble pas cependant avoir été employée pour l'analyse de l'eau de mer, comme on peut le constater par une revue de la littérature scientifique.

La détermination de l'oxygène dissous dans l'eau de mer est d'une importance capitale, tant au point de vue biologique (étude du métabolisme des animaux, des plantes et des micro-organismes vivant dans la mer), qu'au point de vue océanographie physique; dans la détermination des courants marins l'oxygène dissous est maintenant considéré comme facteur aussi important que la salinité et la température. Il y avait donc intérêt à étudier en détail l'application de la polarographie à l'analyse de l'oxygène dissous dans l'eau de mer et les eaux saumâtres afin d'en déterminer la sensibilité et les limites d'emploi. De plus, comme l'électrolyse avec micro-électrode de platine fixe ou tournante peut être adaptée à l'analyse de l'oxygène dissous, une étude de ces électrodes a également été faite.

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Contribution des Laboratoires de Chimie physique de l'Université Laval, Québec, et de la Station biologique du Saint-Laurent.

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Appareils et Méthode de Mesure

Les déterminations ont été faites au moyen du Polarographe de Heyrovský fabriqué par E. H. Sargent. Dans la plupart des cas les courbes courant–voltage ont été enregistrées photographiquement. La sensibilité du galvanomètre était de 0.0032 μamp. par mm. Avec la cellule de Heyrovský le temps de chute des gouttes de mercure était de l'ordre de 4.7 sec. et la constante géométrique du tube capillaire, $m^{2/3}\ t^{1/6}$, avait une valeur de 0.00175 $g.^{2/3}\ sec.^{1/6}$. Dans quelques cas l'anode de mercure a été remplacée par une électrode au calomel à joint rodé, du type Leeds et Northrup, plongeant directement dans le liquide à analyser. Toutefois les potentiels donnés ici ont tous été calculés par rapport à l'électrode au calomel saturée. Pour les analyses avec micro-électrodes de platine la cathode consistait en un fil de platine brillant de 0.5 mm. de diamètre et de 2.8 et 3.2 mm. de longueur pour les électrodes fixe et tournante respectivement; ce fil était soudé à un tube de verre ordinaire rempli de mercure pour faire le contact électrique. L'électrode était soit fixée au bouchon de la cellule (le fil de platine étant en position verticale), soit reliée par des engrenages à l'arbre d'un moteur électrique qui lui communiquait une vitesse de rotation très uniforme de 593 tours par minute (et alors le fil de platine était soudé latéralement). Des montages semblables ont été décrits en détail par Laitinen et Kolthoff (6).

L'eau de mer qui a servi à ces mesures provenait de la baie des Chaleurs dans le golfe Saint-Laurent où elle avait été recueillie à une profondeur de 100 pieds; elle était tamponnée naturellement à un pH de 7.2. Une série de mesures furent également faites sur l'eau du fleuve Saint-Laurent récoltée en face de Québec. L'eau de mer fut employée telle quelle ou diluée avec de l'eau distillée ou encore avec de l'eau du fleuve, tel qu'indiqué. Dans chaque cas la teneur en oxygène fut contrôlée par une analyse chimique (méthode de Winkler) effectuée sur un échantillon de 140 cc. Pour faire varier la quantité d'oxygène dissous dans l'eau de mer il suffisait d'y faire barbotter de l'azote pendant un temps plus ou moins long. Sauf indication contraire les mesures furent faites à la température de la chambre, 22° ± 1.°C.

I. Electrode à gouttes de mercure

Pour cette série de mesures la sensibilité du galvanomètre n'était que le dixième de la valeur indiquée plus haut; soit 0.032 μamp. par mm. Ainsi que le démontre le graphique de la Fig. 1, l'intensité du courant de diffusion à un potentiel de −0.29 v., est rigoureusement proportionnelle au contenu en oxygène de l'eau de mer. Par exemple une eau contenant 5.0 cc. d'oxygène dissous par litre avait un courant de diffusion égal à 2.78 μamp. En tenant compte du courant résiduel, qui atteignait une valeur de 0.07 μamp. à −0.29 v., la précision de l'analyse polarographique est donc de 0.02 cc. par litre, soit environ le double de celle de l'analyse chimique par la méthode de Winkler. Quant à la sensibilité, c'est-à-dire la plus petite quantité d'oxygène qui puisse être décelée, elle est de 0.02 cc. par litre. Il est à remarquer que l'électrolyse des traces de cuivre contenues dans l'eau de mer (de l'ordre de

2×10^{-6} mole par litre) contribue pour une bonne part au courant résiduel: soit environ 0.01 μamp. Afin d'économiser l'eau de mer certaines mesures

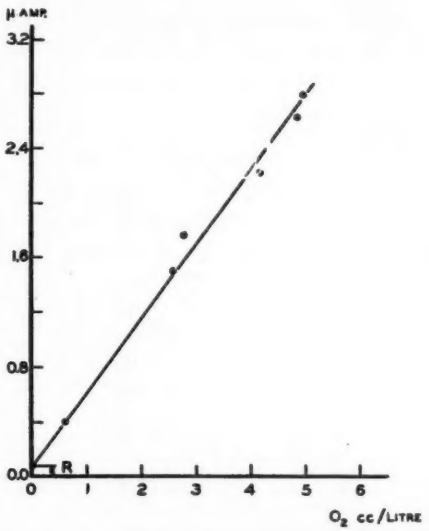

FIG. 1. *Intensité du courant limitatif en fonction de la concentration de l'oxygène dissous dans l'eau de mer, avec la cathode à gouttes de mercure. R: courant résiduel.*

ont été faites avec des solutions de chlorure de potassium de salinité équivalente, soit 0.5 *N*. En effet les résultats de plusieurs mesures comparatives ont démontré que cette substitution était légitime pour toutes fins pratiques (cf. Tableau I).

TABLEAU I

CARACTÉRISTIQUES DE L'EAU DE MER ET D'UNE SOLUTION DE CHLORURE DE POTASSIUM DANS LA RÉDUCTION POLAROGRAPHIQUE DE L'OXYGÈNE

Solution	Potentiel de demi-onde, v.		Courant résiduel à 0.94 v., μamp.
	Première onde	Deuxième onde	
Eau de mer	−0.05	−0.84	0.252
KCl, 0.5 *N*	−0.07	−0.82	0.228

Comme il était à prévoir, il y a assez de sels en solution dans l'eau de mer pour supprimer tout maximum sur la courbe courant–voltage. Ce n'est qu'à une dilution donnant une normalité en chlorure voisine de 0.1 *N*, ou, en d'autres termes, à une salinité de 7.0°/$_{oo}$* qu'un maximum apparaît.

** Par convention on exprime la salinité de l'eau de mer en parties par mille, °/$_{oo}$. L'eau qui a été employée pour ce travail avait une salinité de 30.0 °/$_{oo}$.*

La hauteur de ce maximum croît rapidement à mesure que la dilution augmente, jusqu'à environ 0.01 *N*, après quoi elle diminue. Varasova (4, p. 307) et Kolthoff (5) avaient déjà observé ce phénomène. Les résultats du Tableau II donnent une idée de l'intensité relative de ces maxima et révèlent en même temps une légère différence entre le comportement de l'eau de mer et celui des solutions de chlorure de potassium. En effet, à salinité équi-

TABLEAU II

INTENSITÉ DU MAXIMUM EN FONCTION DE LA CONCENTRATION DE L'ÉLECTROLYTE FORT

Normalité de la solution en ion chlorure	Courant maximum / Courant de diffusion	
	Eau de mer	KCl
0.25 (pas de maximum)	1	1
0.1	1.05	1.5
0.01	9.6	11.1
0.001	5.9	9.2

valente, les maxima dans l'eau de mer diluée sont toujours moins prononcés que dans les solutions salines. Cette différence est attribuable au fait que l'eau de mer contient des anions polyvalents et des substances organiques très actives pour l'élimination des maxima.

Ainsi que le montrent les données du Tableau III, la dilution de l'eau de mer au-delà d'une certaine limite a encore pour effet de déplacer le potentiel au maximum ($E_{max.}$) et même, quoique dans des proportions beaucoup moindres, le potentiel de demi-onde ($E_{1/2}$), c'est-à-dire le point d'inflexion sur la courbe courant–voltage. Pour pouvoir observer ce dernier aux grandes dilutions il faut d'abord ajouter quelques gouttes de méthyle rouge afin de supprimer tout maximum.

TABLEAU III

VARIATION DES POTENTIELS AU MAXIMUM ET DE DEMI-ONDE AVEC LA NORMALITÉ DE LA SOLUTION

N	$E_{1/2}$, v.		$E_{max.}$, v.
	Première onde	Deuxième onde	
0.475	−0.05	−0.84	—
0.25	−0.06	−0.84	—
0.1	−0.06	−0.85	−0.27
0.01	−0.07	−0.86	−0.41
0.001	−0.08	−0.93	−0.57

Kolthoff (4, p. 89) a fait remarquer que lors de la réduction d'une particule non chargée, comme c'est ici le cas pour l'oxygène, le courant limitatif est uniquement un courant de diffusion, et par conséquent, son intensité doit

être indépendante de la concentration de l'électrolyte fort. Les résultats contenus dans le Tableau IV constituent ce qui semble être la première vérification expérimentale de cette proposition.

TABLEAU IV

VALEUR DU COURANT LIMITATIF EN FONCTION DE LA CONCENTRATION DE L'ÉLECTROLYTE FORT

N	O_2	i_l	O_2/i_l
0.0001	4.90	2.50 *a*	1.97
0.0003	5.15	2.50 *a*	2.06
0.001	5.1	2.50 *a*	2.04
0.001	5.2	2.60	1.99
0.01	5.5	2.86	1.92
0.1	5.9	2.96	1.98
0.25	5.6	2.86	1.96
0.475	4.9	2.55	1.92
0.475	2.6	1.40	1.86
0.475	0.6	0.31	1.94
		Moyenne:	1.95 ± 0.1

N: *normalité de la solution en ion chlorure.*
O_2: *concentration de l'oxygène dissous en cc. par litre.*
i_l : *courant limitatif en μamp.*
a : *mesures faites dans des solutions de chlorure de potassium.*

II. Micro-électrodes de platine

Laitinen et Kolthoff (6) ont fait des déterminations d'oxygène avec des micro-électrodes de platine. Dans le cas de l'électrode fixe, il faut attendre deux ou trois minutes avant de faire la lecture du courant limitatif afin de permettre à l'équilibre de s'établir. Contrairement à ce qu'a trouvé Kolthoff pour des solutions tampons, aucun dégagement d'hydrogène n'a été observé dans l'eau de mer (pH 7.2) à −0.74 v. Il a même été possible d'observer, dans l'eau de mer non diluée, le commencement du deuxième palier à −0.94 v. (Fig. 2). La Fig. 3 démontre que l'intensité du courant de diffusion est proportionnelle à la concentration de l'oxygène dissous. Ces mesures ont été faites à −0.61 v. dans une solution de chlorure de potassium décinormale. Le courant résiduel étant de 0.03 μamp., on voit que la sensibilité, de même que la précision, est de 0.02 cc. par litre. Comme pour l'électrode à gouttes de mercure, il existe ici des variations du potentiel de demi-onde en fonction de la concentration de l'électrolyte fort.

Avec l'électrode tournante il a été impossible d'observer le deuxième palier sur la courbe courant–voltage même dans l'eau de mer non diluée. En effet, la courbe de réduction de l'oxygène devient très allongée (Fig. 2) et finit même par disparaître à des salinités voisines de 7 °/₀₀ (0.1 *N*). Avec cette électrode, cependant, le courant de diffusion, et par conséquent aussi le courant résiduel, sont beaucoup plus grands que dans le cas de l'électrode fixe. Il en résulte que la précision des dosages avec l'électrode tournante est quarante fois plus grande qu'avec la cathode à gouttes de mercure.

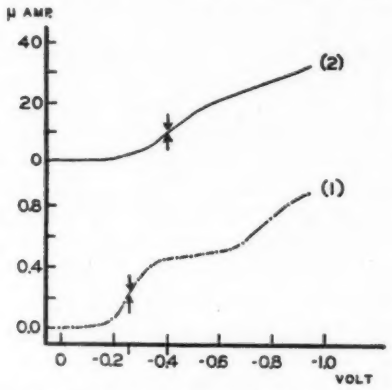

FIG. 2. *Courbes courant–voltage, pour les micro-électrodes de platine fixe (1) et tournante (2).*

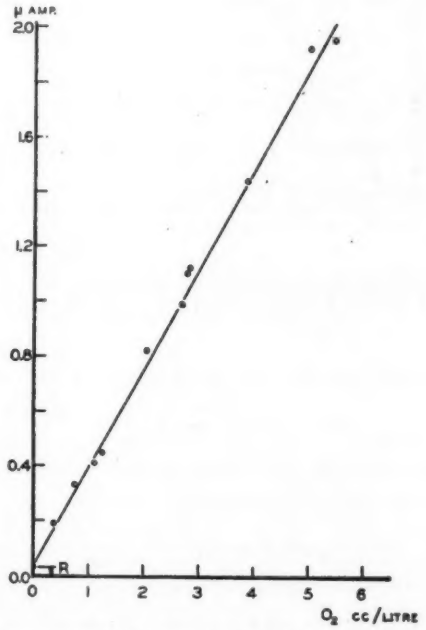

FIG 3. *Intensité du courant limitatif en fonction de la concentration de l'oxygène dans une solution de chlorure de potassium 0.1 N, avec la cathode de platine fixe. R: courant résiduel.*

D'autre part, les résultats représentés sur la Fig. 4 sembleraient indiquer que le courant de diffusion cesse d'être proportionnel à la teneur en oxygène aux concentrations inférieures à 0.5 cc. par litre. Un tel phénomène est difficile à expliquer de prime abord et, comme il ne se rattache qu'indirecte-

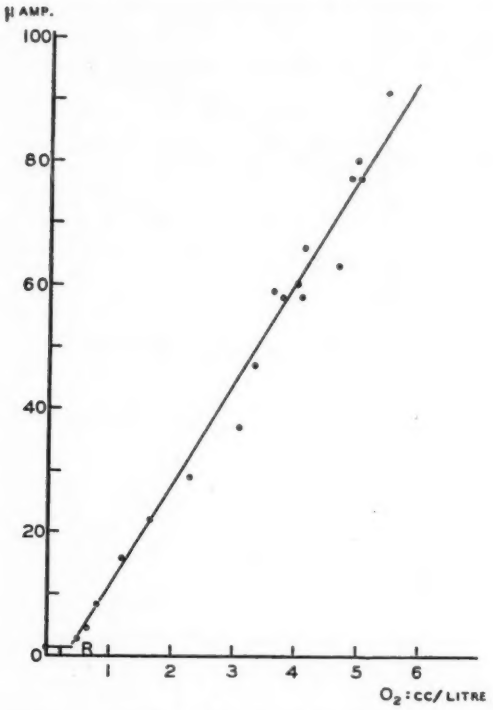

FIG. 4. *Intensité du courant limitatif en fonction de la concentration de l'oxygène dans une solution de chlorure de potassium 0.1 N, avec la cathode de platine tournante. R : courant résiduel.*

ment au présent travail, il a été jugé préférable d'en faire l'objet d'une étude subséquente.

TABLEAU V

RÉSUMÉ DES RÉSULTATS OBTENUS AVEC LES DIFFÉRENTES CATHODES

——	Gouttes de mercure	Electrode de platine	
		fixe	tournante
Sensibilité, cc./litre	0.02	0.02	0.30
$\frac{\text{cc./litre d'oxygène}}{\mu\text{amp.}}$	1.84	2.78	0.05
Courant résiduel, μamp.	0.07	0.03	1.3
$E_{1/2}$, v.	−0.05	−0.25	−0.40

$E_{1/2}$: *dans l'eau de mer (0.475 N).*

Conclusions

La détermination polarographique de l'oxygène dissous dans l'eau de mer peut se faire avec la cathode à gouttes de mercure et les cathodes de platine fixe et tournante aux potentiels de -0.29, -0.61, et -0.91 v. respectivement. La précision et la sensibilité des deux premières cathodes sont environ le double de celles de la méthode de Winkler; l'électrode tournante a une précision encore beaucoup plus grande. Dans l'eau de mer de salinité voisine de 30 °/₀₀, il y a suffisamment d'électrolyte pour éliminer tout maximum; dans les eaux saumâtres, les maxima apparaissent seulement à une salinité de 7 °/₀₀. Le potentiel de demi-onde varie quelque peu aux faibles salinités mais l'intensité du courant limitatif (hauteur du palier dans les courbes courant–voltage) est indépendante de la concentration de l'électrolyte.

Remerciements

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α-CYANO-β-ARYLACRYLIC ACIDS[1]

By C. Y. Hopkins[2], Mary Chisholm,[2] and Ruth Michael[2]

Abstract

α-Cyano-β-arylacrylic acids have been prepared by condensing the following aldehydes with sodium cyanoacetate: 1-naphthaldehyde, 2,3-dimethoxybenzaldehyde, 3,4-diethoxybenzaldehyde, 6-chloropiperonal, 4-isopropylbenzaldehyde, 2-acetoxy-3-methoxybenzaldehyde, 2-hydroxy-3-methoxybenzaldehyde. The last-named gave α-cyano-β (2-hydroxy-3-methoxyphenyl) acrylic acid, which was readily converted to 8-methoxycoumarin-3-carboxylic acid.

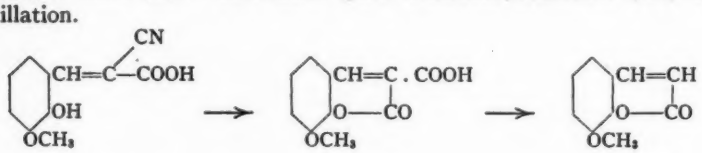
Introduction

It has been shown that certain arylacetic acids may be prepared by the action of sodium hypochlorite on the corresponding α-cyano-β-arylacrylic acids (2). An investigation of the mechanism of this reaction has been undertaken. The present paper describes the preparation of the α-cyano-β-arylacrylic acids.

The condensation of aromatic aldehydes with sodium cyanoacetate was carried out by the general method of Lapworth and McRae (1). Slight variations in the procedure are required, depending on the aldehyde used.

Unless close control of conditions is maintained, a gummy product is obtained which is difficult to purify. This type of product appears to be associated with some hydrolysis of the CN group since free ammonia is formed along with the gummy product.

2-Hydroxy-3-methoxybenzaldehyde condensed readily with sodium cyanoacetate but the product (I) could not be recrystallized without change. It was converted in hot acidified alcohol solution to 8-methoxycoumarin-3-carboxylic acid (II). The latter gave 8-methoxycoumarin (III) by dry distillation.



(I) (II) (III)

Experimental

*α-Cyano-β-(1-naphthyl)acrylic Acid**

1-Naphthaldehyde was prepared from 1-chloromethylnaphthalene by the method of Rupe and Brentano (5). It was purified by means of the bisulphite compound and distilled *in vacuo*.

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Contribution from the Division of Chemistry, National Research Laboratories, Ottawa, Canada. Issued as NRC No. 1260.

[2] *Chemist.*

* *With John Harris.*

A neutral solution of sodium cyanoacetate (total volume, 125 cc.) was prepared from 30 gm. of monochloroacetic acid (1).

A mixture of 28 gm. of 1-naphthaldehyde, 25 cc. of ethanol, and 20 cc. of 20% sodium hydroxide solution was added to the cyanoacetate solution and the mixture was shaken thoroughly. The sodium salt of α-cyano-β-(1-naphthyl)acrylic acid was precipitated. After standing for several hours, the mixture was acidified strongly with hydrochloric acid and the yellow product was filtered off, washed with water and with benzene. The yield was almost theoretical. After recrystallization from acetic acid or ethanol, the substance melted at 225° C.* Found: N, 6.26%; equivalent, 225. Calc. for $C_{14}H_9O_2N$; N, 6.28%; equivalent, 223.

α-Cyano-β-(1-naphthyl)acrylic acid crystallizes from acetic acid in yellow needles. It is readily soluble in acetone and chloroform, moderately soluble in ethanol and acetic acid, difficultly soluble in benzene, and insoluble in water and hexane.

Ethyl α-Cyano-β-(1-naphthyl)acrylate

The ester was prepared from the foregoing acid by refluxing with absolute alcohol and sulphuric acid. It was recrystallized from ethanol, from which it separated in pale yellow needles, m.p. 81 to 82° C. Found: N, 5.53, 5.68%. Calc. for $C_{16}H_{13}O_2N$: N, 5.58%.

It is readily soluble in hexane, benzene, and chloroform, moderately soluble in ethanol, and insoluble in water.

α-Cyano-β-(2,3-dimethoxyphenyl)acrylic Acid

2,3-Dimethoxybenzaldehyde gave this acid in good yield by the procedure described above for 1-naphthaldehyde. After recrystallization from toluene, the acid melted at 222° C. It is soluble in ethanol, slightly soluble in benzene and insoluble in hexane. The ethyl ester was made by heating the acid with absolute alcohol and sulphuric acid. It is soluble in ethanol, chloroform, and benzene, but insoluble in hexane. The melting points of the acid and ester correspond closely to those reported by Perkin and Robinson (4).

α-Cyano-β-(3,4-diethoxyphenyl)acrylic Acid

The condensation of 3,4-diethoxybenzaldehyde was carried out in the presence of alkali as described above, with suitable cooling to prevent the temperature from rising above 25° C. The yield was almost theoretical. The acid was recrystallized from benzene. It forms bright yellow felted needles, m.p. 178 to 179°. Found: N, 5.46%. Calc. for $C_{14}H_{15}O_4N$: N, 5.36%. The substance is very soluble in ethanol and chloroform, moderately soluble in benzene, and insoluble in hexane.

α-Cyano-β-(6-chloropiperonyl)acrylic Acid

Piperonal was chlorinated by means of sulphuryl chloride (6). The chloroaldehyde was dissolved in four times its weight of ethanol and the condensation was carried out in the presence of alkali in the usual way. The product

* *Melting points are corrected.*

was purified by recrystallization from ethanol, forming deep yellow needles, m.p. 280° C. It is soluble in acetic acid, moderately soluble in acetone, slightly soluble in benzene, and substantially insoluble in chloroform and ether. Found: N, 5.44, 5.43%. Calc. for $C_{11}H_6O_4NCl$: N, 5.57%.

α-Cyano-β-(4-isopropylphenyl)acrylic Acid

4-Isopropylbenzaldehyde was condensed with sodium cyanoacetate in the presence of aniline hydrochloride. A solution of 8.8 gm. of the aldehyde in 20 cc. of ethanol was added to a solution of sodium cyanoacetate made from 10 gm. of monochloroacetic acid, to which had been added enough hydrochloric acid to liberate half of the cyanoacetic acid, and containing 5 gm. of aniline hydrochloride. The mixture was warmed to 70° C. and shaken continuously until a solid product had formed. On adding excess hydrochloric acid, an almost quantitative yield of the desired acid was obtained. The product was recrystallized from toluene, forming white needles, m.p. 156 to 157°. Found: N, 6.36%. Calc. for $C_{13}H_{13}O_2N$: N, 6.51%. It is soluble in ethanol, slightly soluble in chloroform, benzene and acetic acid, and insoluble in hexane.

α-Cyano-β-(2-acetoxy-3-methoxyphenyl)acrylic Acid

The acid was prepared from 2-acetoxy-3-methoxybenzaldehyde by the aniline hydrochloride method described above. It was recrystallized several times from toluene, forming small nearly white needles, m.p. 180 to 181° C. Found: N, 5.18%. Calc. for $C_{13}H_{11}O_5N$: N, 5.36%. The substance is soluble in ethanol, acetone, and ethyl acetate, slightly soluble in chloroform and benzene, and insoluble in hexane.

α-Cyano-β-(2-hydroxy-3-methoxyphenyl)acrylic Acid

The condensation of 2-hydroxy-3-methoxybenzaldehyde was carried out by the aniline hydrochloride method. Just sufficient hydrochloric acid was added to liberate the acrylic acid. An excess of mineral acid is detrimental. The product was washed with water, dried, and washed with toluene. It could not be purified by recrystallization but was identified as the acetyl derivative. The latter was prepared by means of acetic anhydride and pyridine. It melted at 180 to 181°C., alone and in admixture with α-cyano-β-(2-acetoxy-3-methoxyphenyl) acrylic acid.

The hydroxy acid forms pale yellow crystals. It melts at about 140 to 160°C. with much effervescence. It is soluble in alcohol but insoluble in benzene and chloroform. The sodium salt is readily soluble in water and gives a deep yellow solution.

When an attempt was made to recrystallize the acid from alcohol, it was quickly transformed in the hot solvent to 8-methoxycoumarin-3-carboxylic acid. The reaction goes to completion in the presence of mineral acid. A mixture of 2 gm. of α-cyano-β-(2-hydroxy-3-methoxyphenyl)acrylic acid, 25 cc. of ethanol, and 1 cc. of concentrated hydrochloric acid was boiled for 10 min. On cooling, 8-methoxycoumarin-3-carboxylic acid crystallized out. After recrystallization from ethanol, it melted at 218° C. Perkin and Robinson (4), who prepared this acid in another way, report the melting point as

210°C. (uncorr.). Found: C, 59.8, 60.1; H, 3.51, 3.77%. Calc. for $C_{11}H_8O_5$: C, 60.0, H, 3.63%. Nitrogen was absent.

The substance crystallizes from ethanol in lemon-yellow needles. It is soluble in chloroform, moderately soluble in ethanol, and almost insoluble in acetone and toluene. When heated at 300 to 350° C., the coumarin acid gave a solid distillate in the form of white crystals, m.p. 89 to 90° after recrystallizing from water. 8-Methoxycoumarin melts at 89° (uncorr.) (3 p. 495).

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A SYNTHESIS OF PHENYLALANINE AND TYROSINE[1]

By Roger Gaudry[2]

Abstract

Acrylonitrile and diazonium chlorides are condensed to give α-chloro-β-arylpropionitriles which can be converted to α-amino acids. Phenylalanine is obtained by transforming α-chloro-β-phenylpropionitrile into the amide with sulphuric acid, hydrolysing to the acid with hydrochloric acid and aminating with ammonia and ammonium carbonate. Tyrosine is obtained from α-chloro-β-*p*-methoxyphenylpropionitrile by direct hydrolysis with mixed acetic and hydrochloric acids and ammonolysis with ammonia and ammonium carbonate.

Introduction

The various syntheses of *dl*-phenylalanine and *dl*-tyrosine have been recently reviewed (6, 7). The best available method for the preparation of *dl*-phenylalanine is probably that of Marvel (5), but there seems to be no really satisfactory method for the preparation of *dl*-tyrosine.

Koelsch (4) obtained α-chloro-β-phenylpropionitrile, with a 34% yield, by condensing acrylonitrile and phenyldiazonium chloride. This suggested to the writer that this compound might be used with advantage for the synthesis of *dl*-phenylalanine and that the method, if successful, might be applied to the synthesis of *dl*-tyrosine.

The writer has already published elsewhere (3) a brief note describing only the principle of the method, as applied to the synthesis of *dl*-phenylalanine. This paper is intended to describe the experimental details of the synthesis, as well as the application of the same method to the synthesis of *dl*-tyrosine. The writer prepared α-chloro-β-phenylpropionitrile with an average yield of 25%, and transformed it into the amide by dissolving it in concentrated sulphuric acid and letting it stand overnight. The amide was hydrolysed to the corresponding acid by boiling with hydrochloric acid. The α-chloro acid was aminated in presence of concentrated ammonia and ammonium carbonate according to Cheronis and Spitzmueller (1), and *dl*-phenylalanine was obtained with a yield of 17% from α-chloro-β-phenylpropionitrile. Only traces of amino acid were obtained when this nitrile was hydrolysed directly by prolonged boiling with concentrated hydrochloric acid and the resulting α-chloro acid aminated in the presence of concentrated ammonia alone.

In order to synthesize *dl*-tyrosine, α-chloro-β-*p*-methoxyphenylpropionitrile was prepared by condensing acrylonitrile and *p*-methoxyphenyldiazonium chloride, in 31% yield. No amide separated when this α-chloro-nitrile was dissolved in concentrated sulphuric acid, left standing overnight, and poured into ice-water. However, when the nitrile was dissolved in a mixture of

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glacial acetic acid and concentrated hydrochloric acid, and hydrolysed directly to α-chloro-β-*p*-hydroxyphenylpropionic acid, which was aminated in presence of concentrated ammonia and ammonium carbonate, crude *dl*-tyrosine was obtained in 20% yield. A fairly large amount of an acid, probably *p*-hydroxycinnamic acid, was obtained when the solution from which *dl*-tyrosine had been filtered was acidified. The *dl*-tyrosine obtained by this method is somewhat difficult to purify, because the α-chloro-nitrile cannot be obtained in a state of purity, since it decomposes slightly during the vacuum distillation, giving off some hydrochloric acid.

Experimental Part

α-Chloro-β-phenylpropionamide

α-Chloro-β-phenylpropionitrile (53 gm.) was dissolved in concentrated sulphuric acid (125 ml.), and the solution was left overnight at room temperature. On pouring on crushed ice, a yellow, crystalline solid separated. It was washed with ice-water (100 ml.) and dried in an evacuated desiccator. Yield, 28.8 gm., 48%. After two recrystallizations from boiling water, it gave long, silky, colourless needles, m.p. 95 to 96°, uncorrected. Calc. for $C_9H_{10}ONCl$: N, 7.63%. Found (Kjeldahl): 7.66%.

dl-Phenylalanine

α-Chloro-β-phenylpropionamide (10 gm.) was boiled under reflux for four hours with 20% hydrochloric acid (100 ml.). The mixture was cooled, extracted with ether, the ether evaporated under a vacuum, and the residue treated with ammonium carbonate (24 gm.), water (10 ml.) and concentrated ammonia (30 ml.) for 60 hr. in a well stoppered, 500 ml. round bottom flask, in a constant temperature oven set at 50° C. The ammonium carbonate and ammonia were evaporated on a boiling water-bath, and the amino acid was recrystallized from hot water. Yield, 3.2 gm., 35.5%. Calc. for $C_9H_{11}O_2N$: N, 8.49%. Found (Kjeldahl): 8.53%. The phenylcarbamyl derivative (2) was prepared from an alkaline solution of the amino acid and phenylisocyanate; m.p. 169 to 170°, uncorrected.

α-Chloro-β-p-methoxyphenylpropionitrile

p-Anisidine (62 gm.) was treated with concentrated hydrochloric acid (100 ml.). The mixture was diazotized in the cold with sodium nitrite (35 gm.) dissolved in water. The diazonium salt was coupled with acrylonitrile (31 ml.) in presence of water (100 ml.), sodium acetate (40 gm.), cupric chloride (15 gm.), and acetone (250 ml.), following closely the method of Koelsch (4) for the preparation of α-chloro-β-arylpropionitriles. The reaction product was extracted with ether, the ether evaporated, and the residue distilled in a vacuum; this treatment yielded α-chloro-β-*p*-methoxyphenylpropionitrile (30 gm.) as a red liquid, b.p. 175 to 177° (8 mm.), with slight decomposition. Calc. for $C_{10}H_{10}ONCl$: Cl, 18.14%. Found: Cl, 15.95%.

dl-Tyrosine

α-Chloro-β-*p*-methoxyphenylpropionitrile (8 gm.) was dissolved in glacial acetic acid (100 ml.), and concentrated hydrochloric acid (100 ml.) was added. The solution was kept for 24 hr. at room temperature, and was then heated under reflux in a boiling water-bath for 16 hr. The solution was evaporated to dryness in a vacuum, and the residue was treated with ammonium carbonate (20 gm.), water (8 ml.), and concentrated ammonia (24 ml.) for 60 hr. at 50° C. in a well stoppered, 500 ml. round bottom flask. The ammonium carbonate and ammonia were evaporated on a boiling water-bath, and the amino acid was precipitated by bringing the solution to pH 5.7 with dilute hydrochloric acid. Yield of crude tyrosine: 1.5 gm., 20%. The amino acid was repeatedly treated with charcoal in boiling water until it separated in characteristic woolly needles, which gave a positive ninhydrin, Millon, and sulphuric acid (Piria) test. Calc. for $C_9H_{11}O_3N$: N, 7.74%. Found (Kjeldahl): N. 7.80%.

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